Reactions Rates: The Arrhenius Equation

Introduction

In this experiment you will determine the Rate Law for the following oxidation-reduction reaction:

$$2\,H^+(aq) + 2\,I^-(aq) + H_2O_2(aq) \rightarrow I_2(aq) + 2\,H_2O(l) \qquad (1)$$

The rate or speed of the reaction is dependent on the concentrations of iodide ion (I^-) and hydrogen peroxide, H_2O_2. (The spectator ions are left off the reaction.) Therefore, we can write the Rate Law (concentration dependence) for the reaction as:

$$\text{Rate} = k\,[I^-]^1\,[H_2O_2]^1 \qquad (2)$$

The temperature dependence of the rate is seen in k – that is, there is a separate value of k for each temperature at which the reaction takes place.

As with a lot of kinetics, the concentration of reactants or products at any instant is difficult to measure directly, so in this lab the rate will be determined indirectly. We have a very handy test for the presence of one of the products, iodine (I_2), namely starch. Starch reacts with iodine to form a blue/black colored complex. Unfortunately as soon as any iodine is produced it will react to make the complex and the solution will turn blue/black instantaneously. Thus, using starch as an indicator by itself would not be of much help. It confirms that some amount of I_2 is being formed, but it tells us nothing about what we are trying to measure - the rate (how long it takes to produce a given quantity of I_2.)

To get around this problem we will introduce a side reaction that will remove the initial I_2 that is produced by our main reaction. This will prevent the solution from turning black long enough so that we can make some time measurements. We will use the following side reaction:

$$I_2(aq) + 2\,S_2O_3^{2-}(aq) \rightarrow 2\,I^-(aq) + S_4O_6^{2-}(aq) \qquad (3)$$

$S_2O_3^{2-}$, thiosulfate ion , reacts with I_2 which prevents the solution from turning blue/black. How will this help? Since we have carefully measured the amount of thiosulfate (a small amount that will run out fairly quickly), we know exactly how much iodine it will take to react with this thiosulfate. Once the small amount thiosulfate has completely reacted, I_2 will start to build up in the solution. As soon as the thiosulfate runs out, I_2 will react with the starch and the solution will turn blue/black. By putting in this "time delay", we can now calculate the rate at which I_2 is being formed.

The rate of reaction is equal to the change in concentration divided by the change in time. The change in time will be the time it takes for the solution to turn dark. We will calculate the change in concentration in I_2 based on the amount of thiosulfate added. Using the known volume and molarity, we can calcuate moles of thiosulfate ($S_2O_3^{2-}$). Based on stoichiometry, we can calculate the moles of I_2 : according to equation **(3)**, 2 moles of thiosulfate react with every 1 mole of I_2, This gives us the change in moles, however for the rate formula we need change in concentration. Divide the moles of I_2 reacted by the total volume to find the change in molarity.

$$\text{Rate} = [\Delta I_2] / \Delta t \qquad (4)$$

Equipment

6 125 or 250 mL Erlenmeyer flasks	Vernier LabQuest	
10 ml and 5 ml Pipettes	Thermometer	
Stop watch or other time keeper	Hot water bath	Ice bath

One bin of chemicals per group that will contain:

0.050 M KI	0.010 M $Na_2S_2O_3$	0.050 M H_2O_2
1.0 M H_2SO_4 ***(Spill : B1)***	1% starch solution	

Disposal: All mixtures Spill/Disposal B1 (down the sink)

Procedure:

1. **Before coming to lab, complete the initial concentration table below showing all work**. Determine initial concentrations using the formula $\mathbf{M_1V_1 = M_2V_2}$. Note that since we are changing temperature, concentration will be kept constant. Copy these values onto the table in the data sheet.

Run	[I^-] (initial)	[H_2O_2] (initial)
1 (Flask 1 + beaker 1)		
2 (Flask 2 + beaker 2)		
3 (Flask 3 + beaker 3)		

2. Clean and mostly dry three Erlenmeyer flasks. Label them 1, 2 and 3.

3. Obtain a bin of chemicals for your group. Use only these chemicals for all of your runs. Use fresh pipettes for each solution. Rinse pipettes twice with the solution that you will be measuring and keep this prepared pipette with the corresponding solution.

4. Add the amounts of the solutions below to prepare each flask. The chemicals must be added in the order listed (top to bottom).

	Iodide Flask #1	Flask #2	Flask #3
0.050 M KI	15.0 mL	15.0 mL	15.0 mL
1% Starch	5.0 mL	5.0 mL	5.0 mL
0.010 M $Na_2S_2O_3$	2.5 mL	2.5 mL	2.5 mL
1 M H_2SO_4	5.0 mL	5.0 mL	5.0 mL

5. Rinse and mostly dry 3 beakers and label as 1, 2 and 3.

6. Prepare the following solutions in clean beakers:

	peroxide flask #1	flask #2	flask #3
0.050 M H_2O_2	15.0 mL	15.0 mL	15.0 mL

7. Get your timer ready. Add the contents of peroxide flask 1 to iodide flask 1. Start the stopwatch as soon as you mix the solutions. Swirl the flask to mix and note the time it takes for the color to change. This is Run 1. Record the temperature of the mixture.

8. Place peroxide Flask #2 and iodide flask #2 in an ice water bath. Let both of them cool until the temperature of both solutions is about 10°C lower than the room temperature. Record the exact temperature, mix the contents of the two flasks and note the time it takes for the color change.

9. Place peroxide Flask #3 and iodide flask #3 in a hot water bath. Let both flasks warm up until the temperature of the two solutions is about 40°C. Record the exact temperature, mix the contents and note the time it takes for the color change.

Disposal: All contents of the reaction flasks may be disposed of into the sink.

Calculations:

The Arrhenius Equation will be used to determine the Temperature dependence. Using the rate law expression determined part 1 (equation 2), calculate k for each temperature.

Using Excel, plot a graph of **ln k** versus **1/T .** If the reaction exhibits Arrhenius behavior, a straight line will be obtained where:

y = m x + b

ln k = − (Ea/R)(1/T) + ln A

where **Ea** is the activation energy of the reaction, **T** is the temperature in **Kelvin**, **k** is the rate constant at this temperature, and **R** is the ideal gas constant (8.314 J/mol K).

- add a best fit line and display the equation.
- Label axes and give the graph a descriptive title
- edit axes such that the data takes up most of the graph
- use the slope of the best fit line to calculate activation energy

Remember to attach the graph and spreadsheet to your lab report.

Name_____

CHM112 Lab –Reactions Rates: The Arrhenius Equation– Grading Rubric

Criteria	Points possible	Points earned
Lab Performance		
Lab work performed correctly. Proper safety procedures followed and waste disposed of correctly. Work space and glassware cleaned up. Participated actively in performing the experiment.	3	
Lab Report		
1. Moles, Molarity and Table	3	
2. Calculations for k	3	
3. Graph and spreadsheet	6	
4. Calculations for E_a	2	
5. Percent error	1	
6. Calculations for k at 90.0 °C	2	
Total	20	

Subject to additional penalties at the discretion of the instructor.

Moles I_2 reacted: _____

Molarity of I_2 reacted [ΔI_2] : _____ (same as values from the Rate Law Lab)

1. Fill in the table. Show all of your work!

Run	Temp (°C)	Initial [I^-] M	Initial [H_2O_2] M	Reaction Time (secs)	Reaction Rate ($\frac{M}{s}$)= ([ΔI_2])/Δt
Cold					
Room temp					
Hot					

2. In the space below show work for the rate constants at each temperature and copy the values into the table below.

3. **Plot a graph of ln (k) vs 1/T in Excel or other graphing program, insert a trendline, and display the equation on the chart. <u>Attach your graph and spreadsheet with the lab report.</u>**

Run	Temperature (in °C)	Temperature (in Kelvin)	1/T in Kelvin	Rate constant 'k' (include proper units)	ln (k)
Cold					
Room temp					
Hot					

Trendline equation: _____

4. Based on your graph, what is the activation energy of the reaction in kJ/mol? Show your calculations below:

5. If the actual value for the activation energy is 53 kJ/mole, using your calculated value for E_a, calculate the percent error.

6. Using your Arrhenius Plot and the best fit equation calculate the value for your rate constant at 65.5 °C.